foot. There is happily a trend toward the use of metric units because most data are collected in these units; but there is also justification for the use of English units because of existing data, especially specifications in such units for air-moving equipment. Conversion factors should always be provided (see the table on page xv). Expressions containing both English and metric units should not be used. The p.p.m. unit (weight to volume) used in water analysis (milligrams per liter) must not be confused with the p.p.m. (volume to volume) unit used in the analysis and expression of atmospheric constituents.

B. VOLUME-PRESSURE-TEMPERATURE RELATIONS

Boyle's law, which states that a perfect gas without change of temperature varies inversely in volume with the pressure to which it is subjected, and Charles' law, which states that at constant pressure the volume of a gas varies directly with the absolute temperature, are usually combined in one operation to convert observed gas or vapor volumes to normal conditions (25° C. and 760 mm. Hg pressure) as follows:

$$\text{Obs. vol.} \times \frac{298}{\text{obs. abs. temp. (°C.)}} \times \frac{\text{obs. bar. pres. (mm. Hg)}}{760} = \text{vol. at normal conditions}$$

This formula may be used, of course, to convert from any observed temperature and pressure to any desired temperature and pressure by substituting the proper values for the normal condition values as given. These industrial hygiene "normal conditions" are not sufficiently different from the "standard air" used in air conditioning and ventilation work (70° F. and 29.92 inches Hg. pressure) to require correction factors in most instances. Likewise these "normals" are sufficiently close to the temperatures and pressures of most occupied spaces that usually no correction is required in industrial hygiene work. The "Standard Temperature and Pressure" used by chemists, however (0° C. and 760 mm.), is sufficiently different (9 per cent) to introduce significant error unless temperature corrections for air volumes are made.

1. Partial Pressures

Each gas in a mixture exerts a pressure in direct proportion to its per cent by volume. In a sample of dry air, for example, at sea level the percentage composition and partial pressures are:

Oxygen: 20.95% × 760 mm. = 159.22 mm. p.p.
Nitrogen, carbon dioxide, and inert gases: 79.05% × 760 mm. = 600.78 mm. p.p.

If, without change in barometric pressure, this air should become humidified to the extent of 50 per cent relative humidity at 25° C., the composition would be altered as follows. Water vapor, 11.88 mm. partial pressure, would appear in the mixture, and since the total pressure does not change, the composition of the atmosphere becomes:

Water vapor = 11.88 mm. p.p.
Oxygen: 20.95% × (760 − 11.88) mm. = 156.73 mm. p.p.
Nitrogen, etc.: 79.05% × (760 − 11.88) mm. = 591.39 mm. p.p.

Should a sample of this 50 per cent humidified air be transported to a higher altitude and opened for analysis, the partial pressures would decrease but the percentage composition would remain the same. At 750 mm. barometric pressure, for instance, this humidified air would have the partial pressure composition, as follows:

$$\text{Water vapor} = \frac{750}{760} \times 11.88 \text{ mm.} = 11.72 \text{ mm. p.p.}$$

Oxygen: 20.95% × (750 − 11.72) mm. = 154.67 mm. p.p.
Nitrogen, etc.: 79.05% × (750 − 11.72) mm. = 583.61 mm. p.p.

Gases are ordinarily reported on a dry basis, but variations in volume due to temperature, pressure, and water vapor changes may be computed by the formula:

$$V_1 = \frac{V_2(P_2 - W_2)(273 + T_1)}{(P_1 - W_1)(273 + T_2)}$$

where V_2 is the observed volume of gas at the observed temperature T_2; V_1 is the calculated volume of gas at temperature T_1; temperatures T_1 and T_2 are in degrees Centigrade; W_2 and W_1 are mm. Hg vapor tension of water at observed and calculated conditions; and P_2 and P_1 mm. Hg are the observed and the calculated barometric pressures, respectively. Air and vapors inhaled into the lungs are rapidly saturated with water vapor at body temperature to approach a partial pressure of 47 mm. Hg and, in computing the partial pressure of any vapor in alveolar air, this partial pressure of water must be considered.

C. ROLE OF SOLUBILITY IN ABSORPTION OF GASES AND VAPORS

1. Solubility Coefficient

Any gas confined in contact with a liquid dissolves at the surface of the liquid and enters the liquid until its partial pressure in the vapor phase above the liquid is in equilibrium with the gas dissolved in the liquid, that is, the rates of its dissolving and evaporating have become equal. The speed with which equilibrium is reached is a function of the solubility and the relationship of contact surface to volume of liquid—the greater the proportion of surface area, and the lower the solubility, the sooner equilibrium will be reached. At the same time the liquid vaporizes into the space above it until its partial pressure in the atmosphere is in equilibrium with its evaporation pressure in the liquid phase. Henry's law, *the*